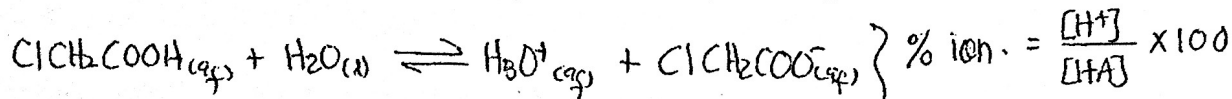


Acid-Base Equilibrium: Weak Acids and Bases (Practice)

1. A 0.100M solution of chloroacetic acid (ClCH_2COOH) is 11.0% ionized. Calculate $[\text{H}^+]$, $[\text{ClCH}_2\text{COO}^-]$, $[\text{ClCH}_2\text{COOH}]$ and K_a for chloroacetic acid.



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{ClCH}_2\text{COO}^-]}{[\text{ClCH}_2\text{COOH}]}$$

$$= \frac{(0.0110)(0.0110)}{(0.100 - 0.0110)}$$

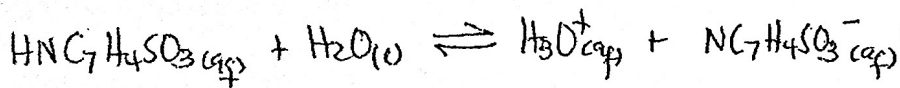
$$K_a = 1.36 \times 10^{-3}$$

$$\% \text{ ion.} = \frac{[\text{H}^+]}{[\text{HA}]} \times 100$$

$$[\text{H}^+] = \frac{(11.0\%)(0.100 \text{ M})}{100} = 0.0110 \text{ M}$$

	[HA]	[H ₃ O ⁺]	[A ⁻]
I	0.100	0	0
C	-0.0110	+0.0110	+0.0110
E	(0.100 - 0.0110)	0.0110	0.0110

2. The $\text{p}K_a$ of saccharin ($\text{HNC}_7\text{H}_4\text{SO}_3$) is 2.32 at 25°C. What is the pH of a 0.10M solution of saccharin?



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{NC}_7\text{H}_4\text{SO}_3^-]}{[\text{HNC}_7\text{H}_4\text{SO}_3]} = 4.8 \times 10^{-3}$$

$$\frac{(x)(x)}{0.10 - x} = 4.8 \times 10^{-3}$$

$$x^2 + 4.8 \times 10^{-3}x - 4.8 \times 10^{-4} = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$x = \frac{-(4.8 \times 10^{-3}) \pm \sqrt{(4.8 \times 10^{-3})^2 - (4)(1)(-4.8 \times 10^{-4})}}{2(1)}$$

$$x = 0.015 ; -0.025$$

$$[\text{H}_3\text{O}^+] = 0.015 \text{ M}$$

$$\text{pH} = -\log(0.015) = 1.81$$

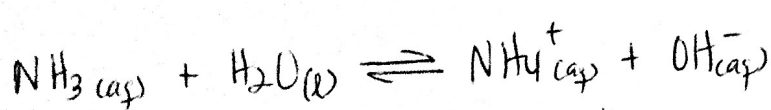
$$K_a = 10^{-\text{p}K_a} = 10^{-2.32} = 4.8 \times 10^{-3}$$

	[HA]	[H ₃ O ⁺]	[A ⁻]
I	0.10	0	0
C	-x	+x	+x
E	0.10 - x	x	x

$$\frac{[\text{HA}]}{K_a} = \frac{0.10}{4.8 \times 10^{-3}} = 20.9 < 400$$

Cannot assume x is small

3. What is the pH of a 0.15 M NH_3 solution? $K_b = 1.8 \times 10^{-5}$



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5}$$

$$\frac{(x)(x)}{0.15 - x} = 1.8 \times 10^{-5}$$

$$\frac{x^2}{0.15} = 1.8 \times 10^{-5} \quad \text{Assume } x \text{ is small}$$

$$x = \sqrt{(0.15)(1.8 \times 10^{-5})} = 1.6 \times 10^{-3} \text{ M} = [\text{OH}^-]$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log(1.6 \times 10^{-3}) = 2.78$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 2.78 = \underline{11.22}$$

	$[\text{NH}_3]$	$[\text{NH}_4^+]$	$[\text{OH}^-]$
I	0.15	0	0
C	-x	+x	+x
E	0.15-x	x	x

$$\frac{[\text{NH}_3]}{K_b} = \frac{0.15}{1.8 \times 10^{-5}} = 8300 > 400$$

Assume x is small

4. Given that the K_b of ammonia (NH_3) is 1.8×10^{-5} and that for hydroxylamine (NH_2OH) is 1.1×10^{-8} , which is the stronger base? Predict which has the strongest conjugate acid. Determine K_a for NH_4^+ and NH_3OH^+ . Was your prediction correct?

NH_3 has a larger K_b . NH_3 is stronger than NH_2OH .

Conjugate acids : NH_3 : NH_4^+

NH_2OH : NH_3OH^+

NH_3OH^+ is the stronger conjugate acid because it's the conjugate acid of the weaker base.

$$K_{a, \text{NH}_4^+} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$K_{a, \text{NH}_3\text{OH}^+} = \frac{1.0 \times 10^{-14}}{1.1 \times 10^{-8}} = 9.1 \times 10^{-7}$$

⇐ Stronger acid because its K_a value is larger